

## Iodine Clock Reaction: Activation Energy Virtual Lab

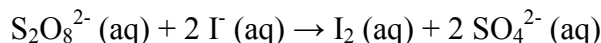
### Introduction

In general, the rate of a reaction increases with increasing temperature. This is consistent with our everyday experiences that cooking food at higher temperatures requires less time than at lower temperatures or that garbage decays faster in the summer than it does in the winter. These observations indicate that the rate constant ( $k$ ) itself depends on temperature. The relationship between the rate constant and temperature is reflected in the Arrhenius equation:

$$k = Ae^{-\frac{E_a}{RT}}$$

where  $k$  is the rate constant,  $A$  is called the Arrhenius pre-exponential constant,  $E_a$  is the activation energy,  $R$  is the gas constant, and  $T$  is the temperature (in units of kelvin). An important point regarding the Arrhenius equation is that the rate constant ( $k$ ) depends on temperature and on the value of the activation energy ( $E_a$ ). The activation energy can be thought of as a barrier to the formation of products and its value is useful in understanding the energetics of the reaction.

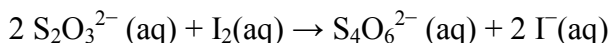
The reaction that we will be studying is known as the iodine clock reaction:



To measure the rate of this reaction we must measure the rate of concentration change of one of the reactants or products. Here, it is convenient to carry out a clock reaction involving the product  $\text{I}_2$ . To do this, you will include (to the reacting  $\text{S}_2\text{O}_8^{2-}$  and  $\text{I}^-$ )

- i) a small (but accurately known) amount of sodium thiosulfate,  $\text{Na}_2\text{S}_2\text{O}_3$ , and
- ii) some starch indicator.

The added  $\text{Na}_2\text{S}_2\text{O}_3$  does not interfere with the rate of above reaction, but it does consume the  $\text{I}_2$  as soon as it is formed (see below):



This reaction is much faster than the previous, so the conversion of  $\text{I}_2$  back to  $\text{I}^-$  is essentially instantaneous. In a typical experiment, known concentrations of  $\text{S}_2\text{O}_8^{2-}$  and  $\text{I}^-$  are mixed with a small amount of  $\text{S}_2\text{O}_3^{2-}$  and starch. As reaction the first proceeds,  $\text{S}_2\text{O}_8^{2-}$  and  $\text{I}^-$  react to produce  $\text{I}_2$ , but nothing appears to be happening because the second reaction is consuming the  $\text{I}_2$  as quickly as it is formed. The important moment is when the  $\text{S}_2\text{O}_3^{2-}$  runs out, for then the  $\text{I}_2$  is no longer consumed (see the second reaction), and the dark blue starch- $\text{I}_2$  complex forms. The stoichiometry of the second reaction indicates that one  $\text{I}_2$  molecule must have been generated for

every 2 ions of  $\text{S}_2\text{O}_3^{2-}$  initially present in the solution. Experimentally, we can determine the value of the activation energy by measuring the value of the rate constant at several measured temperatures of the iodine clock reaction.

### *Procedure*

- 1) Go to the class website and click on the link under “Unit VI-Kinetics and Equilibrium.”
- 2) Make sure that all of the volumes are set to the following:
  - Water—48 mL
  - 0.005 M  $\text{Na}_2\text{S}_2\text{O}_3$  (with starch)—10 mL
  - 1.0 M KI—12 mL
  - 0.1 M  $(\text{NH}_4)_2\text{S}_2\text{O}_8$ —30 mL
- 3) Set the temperature at 25°C.
- 4) Click ‘Start.’ As soon as the last volume is added, click the play button under the Timer.
- 5) Allow the reaction to run until it has turned a dark blue/purple color (will often happen suddenly). Click the pause button under the Timer.
- 6) Record the time in the Data Table.
- 7) Click ‘Reset.’ Repeat steps #4-6.
- 8) Reset everything by clicking ‘Reset.’ Change the temperature to 5°C.
- 9) Repeat steps #4-6 twice. Record each time in the Data Table. Remember to reset everything by clicking ‘Reset.’
- 10) Change the temperature to 15°C, then 35°C, and finally 45°C. Repeat each temperature twice, recording all times in the Data Table.
- 11) Average all of your times for each temperature. Record the average in the appropriate column in the Data Table.

### *Data Table*

| Temperature (°C) | Trial #1 (s) | Trial #2 (s) | Average (s) |
|------------------|--------------|--------------|-------------|
| 5                |              |              |             |
| 15               |              |              |             |
| 25               |              |              |             |
| 35               |              |              |             |
| 45               |              |              |             |

### Calculations

- Calculating the Reaction Rate

The rate can be expressed as:

$$Rate = \frac{-\Delta[S_2O_8^{2-}]}{\Delta t}$$

The  $-\Delta[S_2O_8^{2-}]$  is equal to 0.00025 M, based on the circumstances of this reaction. Calculate the reaction rate for each temperature by dividing this concentration by the average time:

| Temperature (°C) | Average Time (s)* | Reaction Rate (M/s) |
|------------------|-------------------|---------------------|
| 5                |                   |                     |
| 15               |                   |                     |
| 25               |                   |                     |
| 35               |                   |                     |
| 45               |                   |                     |

\*Copied from Data Table

- Calculating the Rate Constant

The rate law for the iodine clock reaction is:

$$Rate = k [I] [S_2O_8^{2-}]$$

The concentration of each reactant is 0.12 M for iodine and 0.03 M for  $S_2O_8^{2-}$ . Using these concentrations and the previously calculated reaction rate, calculate the rate constant (k) for each temperature:

| Temperature (°C) | Rate Constant ( $M^{-1}\cdot s^{-1}$ ) |
|------------------|--|
| 5                |  |
| 15               |  |
| 25               |  |
| 35               |  |
| 45               |  |

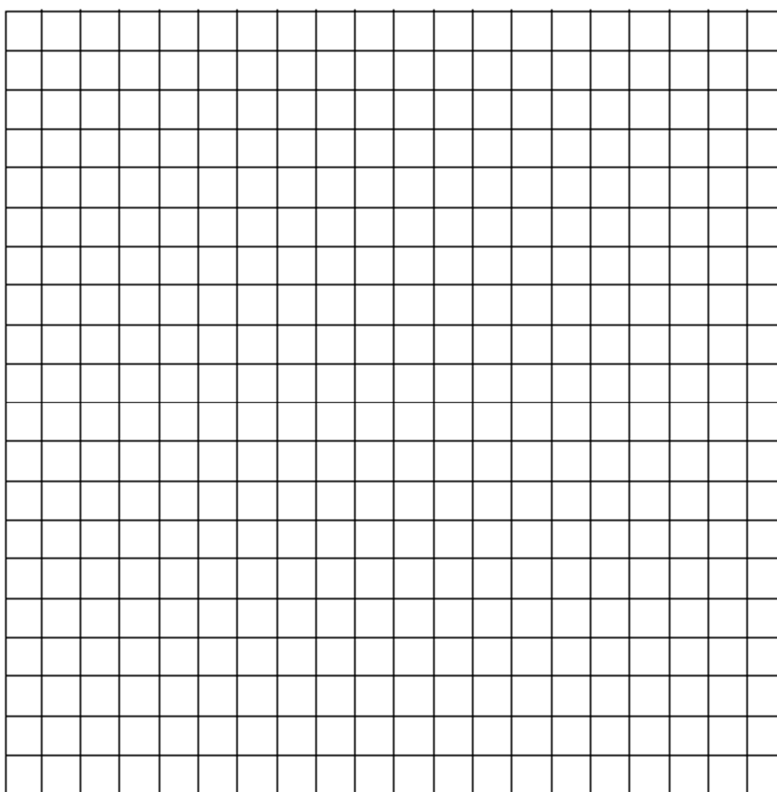
- Determining Activation Energy

The activation energy can be determined graphically by plotting  $\ln k$  versus  $1/T$ . First change the temperatures from Celsius to Kelvin. Then calculate  $\ln k$  and  $1/T$  below:

| Temperature (°C) | Temperature (K) | $1/T$ (K <sup>-1</sup> ) | Rate Constant (M <sup>-1</sup> ·s <sup>-1</sup> )* | $\ln k$ |
|------------------|-----------------|--------------------------|--|---------|
| 5                |                 |                          |  |         |
| 15               |                 |                          |  |         |
| 25               |                 |                          |  |         |
| 35               |                 |                          |  |         |
| 45               |                 |                          |  |         |

\*Copied from previous table

Now graph  $\ln k$  on the y axis and  $1/T$  on the x-axis below. Draw a straight line through the points and determine the slope of the line.



The slope equals  $-E_a/R$ , where  $E_a$  is the activation energy and  $R$  is 8.314 J/mol·K. Calculate the activation energy for this reaction.